

## Chapter 5 Self Assessment, page SG-95

- |      |      |      |       |
|------|------|------|-------|
| 1. D | 4. B | 7. E | 10. B |
| 2. B | 5. E | 8. D | 11. D |
| 3. D | 6. C | 9. A | 12. C |

### Studying Changes in Energy and Heat (5.1), page SG-99

- a. isolated; b. open; c. closed; d. closed; e. isolated (for practical purposes, or closed; no system can be truly isolated except the universe); f. closed
- The first law of thermodynamics states that the energy of the universe is constant. By definition, when studying a system, everything else in the universe is the surroundings. So, if the energy of the universe is to remain constant, any change of energy in a system must be accompanied by an equal and opposite change in the surroundings.

$$\begin{aligned} 3. \quad \Delta T &= T_{\text{final}} - T_{\text{initial}} \\ &= 13.21^{\circ}\text{C} - 22.32^{\circ}\text{C} \\ &= -9.11^{\circ}\text{C} \end{aligned}$$

$$\begin{aligned} Q &= mc\Delta T \\ &= (2.56 \text{ kg})\left(\frac{1000 \text{ g}}{\text{kg}}\right)(0.88 \frac{\text{J}}{\text{g}^{\circ}\text{C}})(-9.11^{\circ}\text{C}) \\ &= -2.0 \times 10^4 \text{ J} \\ &= -20 \text{ kJ} \end{aligned}$$

The water released 20 kJ of heat.

As the concrete cooled, 20 kJ of heat was released.

$$\begin{aligned} 4. \quad Q &= mc\Delta T \\ 2960 \text{ J} &= (25.00 \text{ g})\left(4.19 \frac{\text{J}}{\text{g}^{\circ}\text{C}}\right)(\Delta T) \\ \Delta T &= 28.3^{\circ}\text{C} \\ T &= 38.0^{\circ}\text{C} - 28.3^{\circ}\text{C} = 9.7^{\circ}\text{C} \end{aligned}$$

The initial temperature of the water was  $9.7^{\circ}\text{C}$ .

### Categories of Enthalpy Changes (5.1), page SG-101

- The sealed car is not an isolated system but a closed system, meaning it cannot exchange matter with the surroundings but it can exchange heat with the surroundings. Because the car is not insulated, especially the windows, heat can be transferred to the air inside the car, so that the air inside the car will indeed get very hot.

- If the sample of water were to melt, 23.5 kJ of heat would be absorbed by the system because  $\Delta H^{\circ}_{\text{melt}} = -\Delta H^{\circ}_{\text{free}}$ .
- a. enthalpy of solution,  $\Delta H_{\text{solution}}$   
b. endothermic  
c. The bonds that form between the solute and solvent have a greater total enthalpy because heat is absorbed from the surroundings as the ammonium nitrate dissolves.

### Using Thermochemical Equations to Determine Enthalpy Changes (5.2), page SG-104

- a.  $\text{Ni(s)} + 4\text{CO(g)} \rightarrow \text{Ni(CO)}_4\text{(g)} + 159.6 \text{ kJ}$   
 $\text{Ni(s)} + 4\text{CO(g)} \rightarrow \text{Ni(CO)}_4\text{(g)} \quad \Delta H_r = -159.6 \text{ kJ}$   
b. "As written" means that the enthalpy change corresponds with stoichiometric amounts of the reactants as shown in the equation. In this case, the enthalpy change given can be expected when 1 mol of nickel reacts with 4 mol of carbon monoxide.  
c.  $-319.2 \text{ kJ}$   
d. The enthalpy of a reaction is directly proportional to the amounts of the substances that react.
- a. The enthalpy diagram should show the formulas for the reactants on a horizontal line above the formulas for the products, also on a horizontal line. An arrow between them should point downward. The arrow should have the label  $\Delta H_r = -972.6 \text{ kJ}$ . the y-axis, at left, should have the label  $H \text{ (kJ)}$ ;  
b. Breaking bonds is an endothermic process and forming bonds is an exothermic process; since the overall reaction is exothermic, the formation of bonds must involve the greater energy change.
- a.  $n = \frac{-9684 \text{ kJ}}{-3228.2 \text{ kJ}} \times 1 \text{ mol} = 3.000 \text{ mol}$   
 $m = 3.000 \text{ mol} \times 122.1 \text{ g/mol} = 366.3 \text{ g}$   
b.  $\Delta H = \frac{25.0 \text{ mol}}{3.5 \text{ mol}} \times -3228.2 \text{ kJ} = 2.3 \times 10^4 \text{ kJ}$

### Calorimetry (5.2), page SG-109

- a. simple calorimeter; b. simple calorimeter; c. bomb calorimeter; d. flame calorimeter
- $Q = C\Delta T = (8.79 \text{ J/}^{\circ}\text{C})(8.35^{\circ}\text{C}) = 73.4 \text{ J}$

3. **a.** You need to know the volume and concentration of the solutions. You need to determine which reactant is present in excess quantity; **b.** You need to know the specific heat capacity of water; **c.** You need to measure the initial temperature of the solutions; **d.** You need to measure the equilibrium temperature of the reaction mixture; **e.** You need to assume that the system is isolated, that any thermal energy exchanged with the calorimeter itself (not counting the water) is negligible, and that the solutions are dilute enough to retain the properties of water.

$$\begin{aligned} 4. Q_{\text{solution}} &= m_{\text{solution}} c_{\text{solution}} \Delta T_{\text{solution}} \\ &= (400.0 \text{ g})(4.19 \text{ J/g}^\circ\text{C})(1.60^\circ\text{C}) \\ &= 2681.6 \text{ J} \\ &= 2.68 \text{ kJ} \end{aligned}$$

$$\begin{aligned} n &= cV_{\text{NH}_3} \\ &= (0.500 \text{ mol/L})(100.0 \text{ mL})(1 \text{ L}/1000 \text{ mL}) \\ &= 0.0500 \text{ mol} \end{aligned}$$

$$\begin{aligned} \Delta H &= n\Delta H_r \\ \Delta H_r &= \frac{\Delta H}{n} \\ &= \frac{-2.68 \text{ kJ}}{0.0500 \text{ mol}} \\ &= -53.6 \text{ kJ/mol} \end{aligned}$$

### Enthalpy Changes and Hess's Law (5.3), page SG-112

1. initial and final conditions; enthalpy changes of its individual steps
2. Reverse equation (1). Reverse equation (2) and multiply by 2.
3. Leave equation (1) as is. Reverse equation (2).

$$\begin{aligned} \Delta H^\circ &= -635.1 \text{ kJ} - 178.3 \text{ kJ} \\ &= 813.4 \text{ kJ} \end{aligned}$$

### Standard Molar Enthalpies of Formation (5.4), page SG-114

1. **a.**  $-393.5 \text{ kJ/mol}$ ; **b.**  $-397.7 \text{ kJ/mol}$ ; **c.**  $0 \text{ kJ/mol}$ ; **d.**  $0 \text{ kJ/mol}$ ; **e.**  $-45.9 \text{ kJ/mol}$ ; **f.**  $+81.6 \text{ kJ/mol}$
2.  $\text{C(s)} + \frac{1}{2}\text{H}_2\text{(g)} + \frac{3}{2}\text{Cl}_2\text{(g)} \rightarrow \text{CHCl}_3\text{(l)} + 134.1 \text{ kJ}$
3.  $\Delta H^\circ_r = \sum(nH^\circ_f \text{ products}) - \sum(nH^\circ_f \text{ reactants})$
4.  $\Delta H^\circ = (1 \text{ mol})(-239.2 \text{ kJ/mol}) - (1 \text{ mol})(-110.5 \text{ kJ/mol}) = -128.7 \text{ kJ}$

### Energy in Ontario (5.4), page SG-117

1. First convert kJ/g to kJ/mol for the heat generated by the furnace:

$$\begin{aligned} Q &= 48.30 \text{ kJ/g} \times 16.05 \text{ g/mol} \\ &= 775.2 \text{ kJ/mol} \end{aligned}$$

Now determine the efficiency:

$$\begin{aligned} \text{Efficiency} &= \frac{775.2 \text{ kJ/mol}}{802.5 \text{ kJ/mol}} \times 100\% \\ &= 96.60\% \end{aligned}$$

$$\begin{aligned} 2. \text{ a. } Q_{\text{water}} &= mc\Delta T \\ &= (1.100 \text{ kg})\left(\frac{1000 \text{ g}}{\text{kg}}\right)(4.19 \frac{\text{J}}{\text{g}^\circ\text{C}})(+13.72^\circ\text{C}) \\ &= +63235 \text{ J} \\ &= +63.24 \text{ kJ} \end{aligned}$$

$$\begin{aligned} Q_{\text{container}} &= mc\Delta T \\ &= (120.4 \text{ g})(0.8967 \frac{\text{J}}{\text{g}^\circ\text{C}})(+13.72^\circ\text{C}) \\ &= +1481 \text{ J} \\ &= +1.481 \text{ kJ} \end{aligned}$$

$$Q_{\text{total}} = 63.24 \text{ kJ} + 1.481 \text{ kJ} = 64.72 \text{ kJ}$$

To calculate the total energy released by burning the marshmallow, factor in the efficiency of the calorimeter.

$$Q = 64.72 \text{ kJ} \div 0.673 = 96.2 \text{ kJ}$$

- b.** nutritional calories (Cal) in marshmallow

$$= \frac{96.2 \text{ kJ}}{4.19 \text{ kJ/Cal}} = 23.0 \text{ Cal}$$

3. **a.** hydroelectricity; **b.** wind power; **c.** natural gas; **d.** nuclear power
4. Sample answer:

#### Risks and Benefits of Using Coal to Generate Electricity

Risks	Benefits
<ul style="list-style-type: none"> <li>• non-renewable</li> <li>• low efficiency</li> <li>• produces greenhouse gases</li> <li>• produces gases that contribute to acid precipitation</li> </ul>	<ul style="list-style-type: none"> <li>• low cost</li> <li>• large available supply of coal</li> <li>• power plants can accommodate changing demands for energy</li> </ul>